



Hyperbaric Measurements Guidance Document

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1 Introduction

This document aims to explain the concepts behind the measurement units which are often used when measuring gases. Many people are familiar with measurement of gas at normal atmospheric pressure. Generally such measurements are 'simplified' as an aid to understanding, but this has the unfortunate effect of complicating the user's understanding of hyperbaric measurements.

It is common for gas concentrations simply to be referred to by their volume ratio expressed simply as a percentage. This is the ratio of the volume of a gas to the total volume occupied by all the components of a gas mixture at the same temperature and pressure.

Less frequently, a weight ratio is used (also expressed as percent), and sometimes a weight per unit volume such as milligrams per cubic metre (mg/m^3).

In this document we will be discussing the volume ratio, or the volumetric percentage. Note that for very small values, this is often expressed in parts per million (ppm). This is purely for convenience, so don't be put off by the term.

One million ppm is the same as 100%, and it just gives another way of describing small gas concentrations, as shown in Table 1.

Table 1: Parts per million and percent conversion

Parts per million (ppm)	Percent (%)
1000000	100
100000	10
10000	1
1000	0.1
100	0.01
10	0.001
1	0.0001

In hyperbaric applications, if we know the volumetric percentage of a gas, this data may not be informative without also knowing the pressure at which the measurement was made. The concept of partial pressure becomes a more useful measurement which can be much more informative with regard to life support conditions.

So please read on and hopefully we can explain the benefits of using either partial pressures or volumetric percentages.

2 The air that we breathe

We have all become quite accustomed to living in the atmosphere of our planet Earth, particularly at sea level and a little higher. The atmosphere has a remarkably 'constant' composition. Remarkable because of the dynamic processes that are going on such as our use of oxygen and our production of carbon dioxide to name just one.

Various sources of information exist to detail the composition of the atmosphere, although they tend to be slightly different albeit at a fairly minute detail. Disagreements about the actual concentrations of the greenhouse gases (carbon dioxide, methane, ozone etc) are part of the reason for the differences.

So in broad terms we can say that the atmosphere consists of the gases listed in Table 2. We've rounded up some of the numbers here, so don't be surprised that the total doesn't add up to exactly 100%

Table 2: Composition of Earth's atmosphere

Name of Gas	Molecular Formula	% vol content
Nitrogen	N ₂	78.08
Oxygen	O ₂	20.948
Argon	Ar	0.93
Carbon dioxide	CO ₂	0.039
Neon	Ne	0.0018
Helium	He	0.00052
Methane	CH ₄	0.00018
Water vapour	H ₂ O	typically between 1% and 4% at Earth's surface
+trace amounts of krypton, hydrogen, nitrous oxide, carbon monoxide, xenon, ozone, nitrogen dioxide, iodine, ammonia and others	Kr, H ₂ , N ₂ O, CO, Xe, O ₃ , NO ₂ , I ₂ , NH ₃	

Water vapour is one of the major players in trying to determine the actual contents of the main gases. The percentages shown refer to a dry atmosphere. Any moisture present will partially displace all of the other gases in proportion. For now just let us worry about the dry atmosphere. (Note – Reference document RM-003 details the effect of moisture on gas composition at different temperatures).

By saying we've become accustomed to the atmosphere, what we mean is that we can live comfortably whilst breathing in these gases. At these levels, we essentially have enough oxygen to support life, not too much carbon dioxide to poison us and also importantly not too much nitrogen. We often say the nitrogen acts as a fairly inert gas, but in fact at higher concentrations (as experienced in air diving) it has a narcotic effect. Hence why deeper forms of diving (eg saturation diving) use helium rather than nitrogen as a background gas.

Information published by safety authorities (for example HSE in the UK or OSHA in the US) specifies safety limits for various pollutants in the atmosphere. These relate to atmospheric pressure, and the limits are given in percentage volume terms. For instance the HSE EH40 document states that the occupational safety limits for carbon dioxide are 5000ppm (for long term 8 hour exposure) or 15000ppm (for short term 15 minutes exposure). This in fact is only true at normal atmospheric pressure. At higher pressures (deeper depths when diving), the human body will only withstand lesser concentrations of carbon dioxide. More of that soon...

3 Partial pressure

Partial pressure is a term often used in hyperbaric applications to quantify gas concentrations. But what do we mean when we talk of partial pressure?

In school you were probably taught Dalton's Law which states that the partial pressure of an ideal gas in a mixture is equal to the pressure it would exert if it occupied the same volume alone at the same temperature. For our purposes here, we will assume that real gases behave the same as ideal gases, which is so nearly true, that we need not concern ourselves about the differences.

So let us take the main gases from Table 2 down as far as carbon dioxide.

Also let us say that atmospheric pressure is 100kPa (to use the SI units), or more commonly (certainly in the UK), we would say 1000 mbar.

Since 100% of the atmosphere exerts a pressure of 1000 mbar, we can create Table 3 in which we state the partial pressures of each of the gases.

We've used 100kPa here because it keeps the numbers nice and easy to handle. In reality atmospheric pressure varies – typically from 95 to 105 kPa (950 to 1050 mbar). In fact, standard atmospheric pressure is defined as 101.325 kPa (1013.25 mbar or 760mm Hg or 14.696 psi depending on which units you care to use). This doesn't really matter in this argument, since what we are interested in is much larger variations in pressure.

Table 3: Partial pressures of gases in atmosphere at 100kPa (1000mbar)

Name of Gas	Molecular Formula	% vol content	Partial Pressure (kPa)	Partial Pressure (mbar)
Nitrogen	N ₂	78.08	78.08	780.8
Oxygen	O ₂	20.948	20.948	209.48
Argon	Ar	0.93	0.93	9.3
Carbon dioxide	CO ₂	0.039	0.039	0.39
Others		0.003	0.003	0.03
Total		100.00	100.00	1000.00

Now look at what happens in Table 4 where we've taken our sample of air, and we've doubled the pressure using the same source of gas. The percentages in the atmosphere (now at 200kPa, 2000 mbar) have remained exactly the same. But the partial pressure of each of the gases has doubled, such that we can see that the total pressure now adds up to the expected 200kPa. So taking oxygen as an example, we can see that the partial pressure has been increased from 209.5 mbar up to 419 mbar by the change in pressure.

Table 4: Partial pressures of gases in atmosphere at 200kPa (2000mbar)

Name of Gas	Molecular Formula	% vol content	Partial Pressure (kPa)	Partial Pressure (mbar)
Nitrogen	N ₂	78.08	156.16	1561.6
Oxygen	O ₂	20.948	41.896	418.96
Argon	Ar	0.93	1.86	18.6
Carbon dioxide	CO ₂	0.039	0.078	0.78
Others		0.003	0.006	0.06
Total		100.00	200	2000

We can say for any gas, that the

$$\text{Partial Pressure (PP}_{\text{gas}}) = V/100 * P_{\text{Total}} \quad [\text{equation 1}]$$

where V = volumetric percent concentration of that gas in the overall sample
and P_{Total} = total pressure of the sample.

As expressed here, V will be a number between 0 and 100, and PP_{gas} and P_{Total} will both be expressed in the same units of pressure as each other. So they could both be kPa, mbar or whatever units with which we choose to measure pressure.

Now, some people don't like to measure gases or refer to their concentration in terms of a pressure. They like using a percentage term. But the straightforward volumetric percentage is not always the most useful of numbers.

Take the situation of a saturation diver at a pressure of 20 bar absolute, and breathing from an atmosphere containing 400 mbar partial pressure of oxygen (400 mbar ppO₂).

From equation 1, we can calculate that the volumetric concentration = $0.4 * 100 / 20 = 2.0\%$

You might be concerned to think that the diver is breathing 2% oxygen. It doesn't sound very much. But in fact the diver is perfectly happy with this level of oxygen. They are actually experiencing almost twice the level of oxygen they would breathe normally at the surface. It is the 20 bar absolute pressure that masks the situation.

So to overcome the confusion, we introduce the term "surface equivalent".

Imagine collecting up all of the molecules of oxygen making up the 400 mbar ppO₂. Having collected them, we will take them to the surface and place them into an imaginary container which initially contains a vacuum. So inside the container we still have 400 mbar ppO₂. We now introduce an inert gas into the container (eg helium or nitrogen – it really doesn't matter – we are still in an imaginary environment). We top up the oxygen with enough inert gas to raise the total pressure to 1000mbar. 600mbar partial pressure of nitrogen (600mbar ppN₂) will suffice. So now we have 400mbar ppO₂ at a total pressure of 1000mbar. Or expressed as a percentage, 40% of our imaginary atmosphere is oxygen. So back down to the diver now at depth, and we can say that his surface equivalent level of oxygen is 40% SEV where SEV stands for surface equivalent value.

Remember we said the diver was breathing about twice the level of oxygen they would breathe at the surface. Well now the numbers fit nicely – 20% (or 20.9%) at the surface, and in this case 40%SEV at depth. So you can see why some people find this a nice easy scale to use – it makes the numbers meaningful.

It does however give rise to one complication worthy of mention. What if we were in a situation breathing pure oxygen (100% O₂) at a pressure of 2 bar absolute? Equation 1 gives us that the partial pressure of oxygen in this case is 2 bar ppO₂ or 2000 mbar ppO₂. Taking these oxygen molecules to the surface to put in our imaginary vacuum container is now more awkward. We can fit in 1000mbar of ppO₂ and give ourselves 100%SEV, but what about the remainder. Well since we've used up half of the oxygen to get to 100%SEV, then the obvious answer is to say that using all the oxygen gives us 200%SEV.

Note that just as 400mbar ppO₂ is the the same as 40%SEV, and now 2000mbar ppO₂ is the same as 200%SEV, we can simply say

$\text{SEV} = \text{PP}_{\text{Gas}}/10$ where PP_{Gas} is the gas partial pressure measured in mbar. Note you can see here that the surface equivalent value, although expressed as a percentage, is directly analogous to the partial pressure. Therefore the units %SEV implies that the measurement is a partial pressure.

4 Choice of partial pressure or percent measurement

So how do we decide whether we should measure gas concentrations as partial pressures or as volumetric percentages?

Measurements made at normal atmospheric pressure tend usually to be done as straightforward volumetric percentages, or as 'parts per million' for gases at low concentrations.

For pressurised applications, the first question to ask is why are we trying to measure the gas?

Take oxygen again as an example, which is usually measured for one of two reasons:

- a) because it is a vital requirement of a life support system for people to breathe
- b) because its presence directly affects the risk of fire

So let us look at each of these in more detail.

4.1 Life support – sustaining life

As we have mentioned earlier in this document, we know that humans are happy to breathe around 200 mbar ppO₂. We can manage with less, but our ability to work hard diminishes the lower the level of oxygen, and eventually we will suffer from hypoxia. This will affect different people at varying degrees. Suffice to say for now that a few very fit and healthy people can climb Everest without the use of oxygen supplies. At the summit of Everest (8848m altitude), atmospheric pressure is approximately 330mbar, hence there will be around 66 mbar ppO₂. To survive unaided at this altitude is said to require an enormous increase in the rate of breathing, and many people would be unable to survive. In fact altitude sickness in less able humans commonly occurs in excess of 2400m, where there will be approximately 150 mbar ppO₂.

At the other end of the scale, exposing ourselves to significantly more than 200mbar ppO₂ will ultimately lead to hyperoxia. This tends to be time related, so breathing oxygen at higher concentrations is generally limited to shorter durations to avoid experiencing problems. Hyperbaric treatment therapies are generally limited to use of pure oxygen at a maximum pressure of 2.8 bar absolute (2800mbar ppO₂), and even then, the patients are given frequent air breaks (about every 20 minutes) to avoid oxygen toxicity problems.

Diving using air is generally limited to 60 metres (7 bar absolute, approximately 1450 mbar ppO₂) to avoid oxygen problems. To go deeper, the oxygen content is reduced below 20%. This is not achieved by increasing the nitrogen content, since this would cause nitrogen narcosis problems. Depending on the scenario, either helium is used to partially replace some of the nitrogen (forming trimix gas), or the nitrogen is completely replaced with helium (forming a heliox mixture).

The actual content of oxygen in a diving bottle would be calculated according to the depth at which it is to be used.

The important factor in all of these applications is that the partial pressure of oxygen experienced by the person remains within safe limits. In terms of oxygen alone, the actual depth is effectively unimportant, other than working out what percentage of gas needs to be blended.

Let us now turn our attention to the carbon dioxide which we produce as a result of our respiration. We said earlier that we typically should avoid breathing 5000ppm of carbon dioxide for longer than 8 hours. Just as with oxygen, it is the partial pressure that affects our bodies. So the 5000ppm at the surface in this case equates to 5000 ppmSEV which can also be expressed as 0.5%SEV or 5 mbar ppCO₂. At the surface, the 8 hour exposure would tend to make us drowsy, which is not the best way to treat our divers. Therefore the 8 hour exposure would tend to be dropped, and the 5 mbar ppCO₂ limit applied as a first alarm threshold. Generally the atmosphere in a diving system would be scrubbed to remove carbon dioxide. Whilst the scrubbers are working at peak efficiency, it is likely that the CO₂ is being maintained at a very low level. It is as the scrubbers lose efficiency and require replenishment that the CO₂ levels will increase fairly quickly to the alarm setpoint. It would be usual best practice to take an immediate action from the alarm, and typically change over the active scrubber units, and replenish the scrubber material in the now exhausted scrubber.

Now to put some numbers to this. In many diving systems, for instance the living chambers in a saturation system, a sample of gas is taken from the chamber and reduced in pressure to be analysed at atmospheric pressure. Such an analyser would typically tell us the carbon dioxide content in parts per million. The alarm setpoint corresponding to the important 5 mbar ppCO₂ level, will vary depending on the pressure in the chamber, as shown in Figure 1. Take an example at 20 bar absolute, we can see that the alarm setpoint is 250ppm. Working backwards, 250ppm represents 0.025% CO₂, and 0.025% of 20 bar abs is 0.005 bar ppCO₂ which is 5 mbar ppCO₂.

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Alternatively, we could mount a carbon dioxide sensor in the chamber itself, and it could give us a direct readout in partial pressure of carbon dioxide.

The advantage of the latter method is that the alarm setpoint can effectively remain fixed at 5mbar ppCO₂ throughout the operation of the system at all pressures. Whereas, with the system measuring ppm CO₂, we need to adjust the alarm setpoint to the appropriate setting for any given depth.

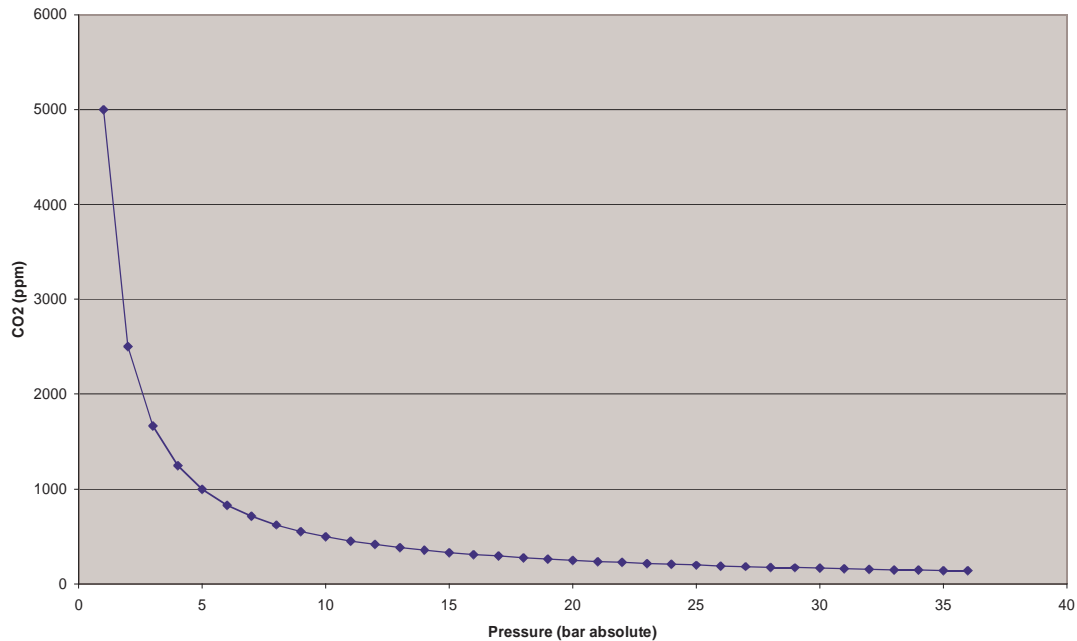


Figure 1: Typical CO₂ (ppm) alarm setting at different pressures

We have an added complication here. Some users have spent many years analysing samples at the surface, and their operating procedures guide them to setting alarm setpoints appropriate to each depth stop in ppm CO₂. When some of these users have taken delivery of newer equipment which is capable of being mounted inside the pressure chamber, they sometimes have a strong preference to retain the measurement in ppm CO₂. Others embrace the simplicity of the partial pressure readout, and the fact that the alarm setpoint is fixed and easily understood.

Analox caters for both groups of users, and makes the option available to display the gas readings in either volumetric format (ppm, % etc), or in partial pressures. Indeed we even try to simplify the task of maintaining alarm setpoints for the user who wishes to use volumetric readouts, and we can arrange that our systems automatically calculate the alarm setpoints according to the pressure as in Figure 1.

4.2 Fire risk

Another application of oxygen sensors is to monitor an atmosphere to minimise the risk of fire. It is generally considered that the risk of fire increases as the percentage volume of oxygen increases (not the partial pressure). The simplest way to think of this is that in normal air, the nitrogen in the air suppresses the ability of oxygen to support a fire. Increasing the pressure, but keeping the oxygen/nitrogen ratio the same (air) does not markedly increase the risk. However any increase in the oxygen/ nitrogen ratio will increase the fire risk.

Situations where this is applied may be where pure oxygen is used e.g. for a fighter aircraft pilot's mouthpiece, or for therapy in a hyperbaric chamber. In either case, the pure oxygen can leak into the cockpit or chamber and enrich the oxygen level. Generally the alarm setpoint will be set to between 23 and 25% oxygen. For guidance in such matters though, refer to the standards for your particular application.

A 1500 mbar or 2000 mbar ppO₂ sensor is adequate to measure 25% oxygen at pressures up to 6 bar abs or 8 bar abs. Beyond this pressure, the sensors would be over-range. We can supply oxygen sensors with higher ranges, but this is not normally required. At these pressures, users would generally start to control the atmosphere to significantly reduce the oxygen content to prevent hyperoxia. As the oxygen percentage decreases, the fire risk also decreases. However, don't forget that pure oxygen may still be used in such a system to maintain the oxygen level as the divers breathe the oxygen levels down. Hence the fire risk remains high, particularly around the oxygen system, hence the particular care paid to oxygen cleanliness in such systems.

Users predominantly interested in fire risk will generally opt to display oxygen in volumetric percent format.

5 Recap

Hopefully, by reading the above, we may have helped you to understand more about partial pressure. So just to check lets refer to some examples, and lets put numbers to them

5.1 Air system

Assume we have a system which can be taken to 10 bar absolute, and that the system is pressurised using air which contains 20.9% oxygen and 400ppm carbon dioxide. Table 5 shows the corresponding gas measurements in percent, mbar and percent surface equivalent. Figure 2 shows the % and %SEV measurements on a graph to clarify that the volumetric percent readings will remain constant, but the %SEV partial pressure readings will increase with pressure.

Table 5: Air system pressurised up to 10 bar absolute

Pressure bar abs	Oxygen			Carbon dioxide			
	%	mbar	%SEV	ppm	%	mbar	%SEV
1	20.9	209	20.9	400	0.04	0.4	0.04
2	20.9	418	41.8	400	0.04	0.8	0.08
3	20.9	627	62.7	400	0.04	1.2	0.12
4	20.9	836	83.6	400	0.04	1.6	0.16
5	20.9	1045	104.5	400	0.04	2	0.2
6	20.9	1254	125.4	400	0.04	2.4	0.24
7	20.9	1463	146.3	400	0.04	2.8	0.28
8	20.9	1672	167.2	400	0.04	3.2	0.32
9	20.9	1881	188.1	400	0.04	3.6	0.36
10	20.9	2090	209	400	0.04	4	0.4

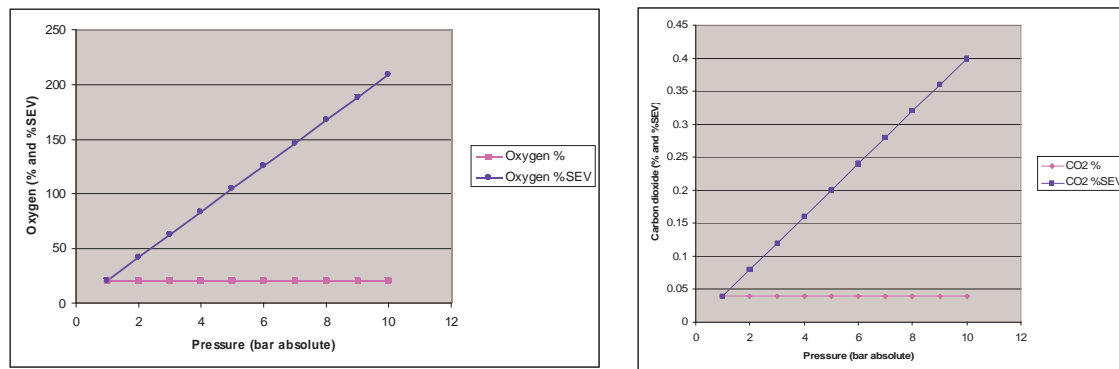


Figure 2: Air system pressurised to 10 bar absolute

5.2 Sat diving system

Assume we have a system which can be taken to 40 bar absolute, and that the system is maintained at an oxygen content of 400 mbar ppO₂. Note in reality, the oxygen level would not be taken to 400mbar ppO₂ until attaining a pressure of around 2 bar absolute. This prevents the percentage of oxygen ever exceeding 20%, and hence minimises risk of fire. Also assume that we wish to maintain a CO₂ alarm setpoint of 5 mbar ppCO₂ (0.5% SEV)

Table 6 shows the corresponding gas measurements in mbar, percent surface equivalent and percent, whilst Figure 3 shows the % and %SEV measurements on a graph to clarify that a volumetric percent reading decreases with pressure, whilst the partial pressure reading in %SEV will remain constant.

Table 6: Saturation system pressurised up to 40 bar absolute

Pressure bar abs	Oxygen			Carbon dioxide			
	mbar	%SEV	%	mbar	%SEV	%	ppm
1	400	40	40.00	5	0.5	0.5000	5000
2	400	40	20.00	5	0.5	0.2500	2500
3	400	40	13.33	5	0.5	0.1667	1667
5	400	40	8.00	5	0.5	0.1000	1000
7	400	40	5.71	5	0.5	0.0714	714
10	400	40	4.00	5	0.5	0.0500	500
15	400	40	2.67	5	0.5	0.0333	333
20	400	40	2.00	5	0.5	0.0250	250
25	400	40	1.60	5	0.5	0.0200	200
30	400	40	1.33	5	0.5	0.0167	167
35	400	40	1.14	5	0.5	0.0143	143
40	400	40	1.00	5	0.5	0.0125	125

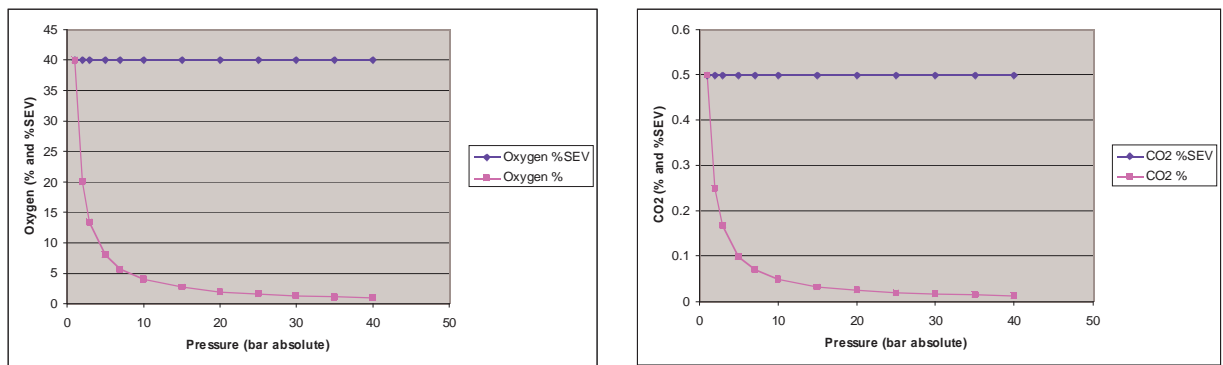


Figure 3: Saturation system pressurised up to 40 bar absolute

6 Choice of measurement units

So by now hopefully we can decide whether a particular system is best suited to either volumetric measurements or to partial pressure measurements. There are then further choices to suit individual customers. Often these are related to the country in which a system will be used since many countries have their own preferred units of measurement.

6.1 Pressure measurement

Throughout this document we've talked of pressure in either kilopascals or bars.

Pascals are after all the SI unit of measure, but very few customers have so far adopted this unit. Kilopascals are ideal for air type systems (1000kPa = 10 Bar), whilst Megapascals tend to be used for saturation systems (40MPa = 40 Bar).

The Bar unit is often used in the UK, whereas some other customers prefer to use Atmospheres or ATA (1 ATA = 1.01325 Bar). This makes sense in that 1 ATA is actually the pressure referred to by the term 'standard atmospheric pressure'.

We mentioned earlier the unit mmHg (millimetre of mercury). These units tend to be used far less frequently than a few years ago, presumably since mercury barometers became more scarce! 1 mmHg was also referred to as 1 Torr by some users, and this gave us $760 \text{ Torr} = 1.01325 \text{ Bar} = 1 \text{ ATA}$.

And finally, since most hyperbaric applications are water based, this gives rise to using a depth unit, which nowadays is generally msw (metres of sea water) or fsw (feet of sea water).

One complication here, is that the actual (or absolute) pressure at a given depth is due to the depth of the water column above it in addition to the atmospheric pressure on the surface.

A depth of zero therefore relates to atmospheric pressure. The factors $10 \text{ msw} = 1 \text{ bar}$ and $32.8083 \text{ fsw} = 1 \text{ bar}$ are often used, such that

1 bar absolute = 0 msw = 0 fsw,
and 2 bar absolute = 10 msw = 32.8083 fsw.

Note these figures are always approximations, because the actual pressure depends on the density of the sea water, which varies both with temperature and salinity around the world.

6.2 Oxygen and carbon dioxide measurement

If we measure partial pressures, then we could use any of the units of absolute pressure. So we could use bars (bar), millibars (mbar), atmospheres (ATA or mATA), kilopascals. And of course we could use the %SEV units we have described in this document. Since carbon dioxide will always be present in lower quantities than oxygen (certainly in life support applications), we also have the ppmSEV units also discussed.

If partial pressures don't suit the application, then we can always revert to the volumetric units of percent or parts per million.